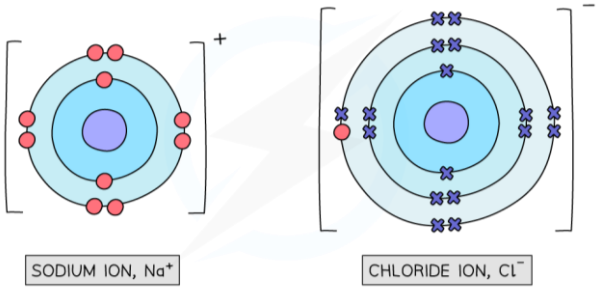
Topic 3: Bonding and Structure

3A: Ionic bonding

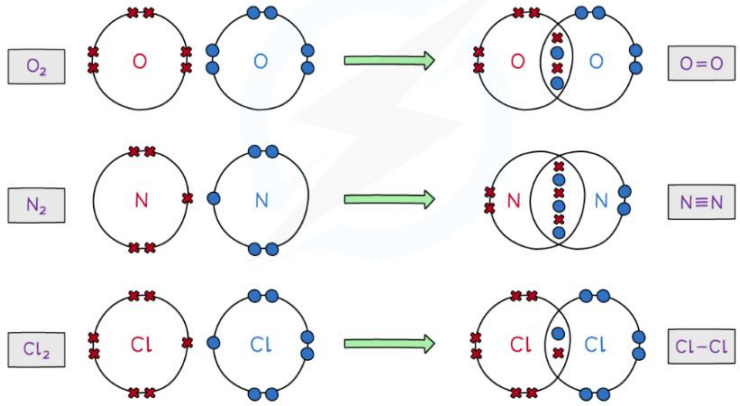
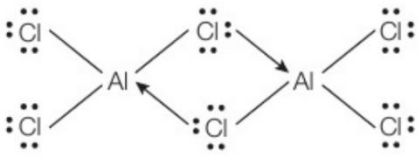
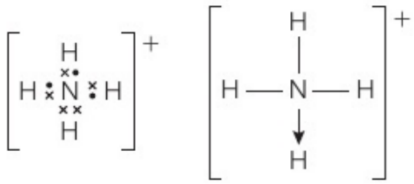
Students will be assessed on their ability to:

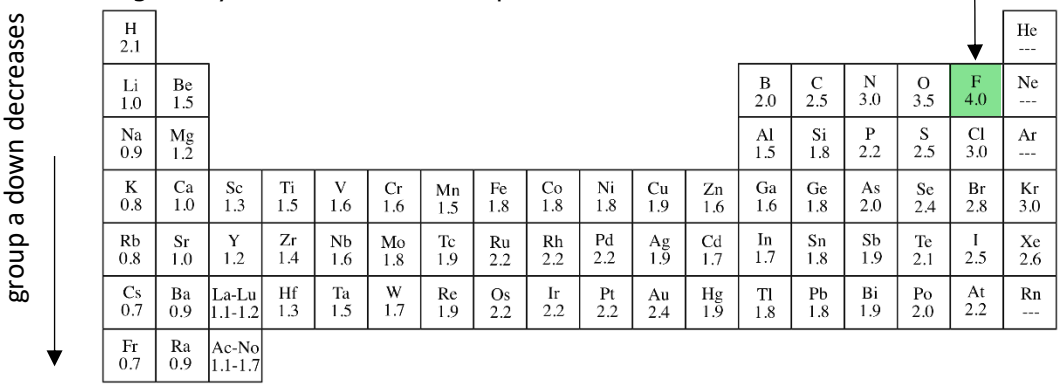
3.1	know and be able to interpret evidence for the existence of ions, limited to physical properties of ionic compounds, electron density maps and the migration of ions
	The behavior of ionic substances in it being able to conduct electricity and undergo electrolysis when either molten or in aqueous solution provides evidence for existence of ions as the positive ions in the solution are attracted towards the negative electrode and the negative ions are attracted towards the positive electrode This can be seen with the electrolysis of green copper(II)chromate(VI) in which the blue Cu^{2+} ions can be seen nearer the negative electrode and the yellow CrO_4^{2-} ions can be seen closer to the positive electrode Electron density maps also provide evidence for the existence of ions
3.2	be able to describe the formation of ions in terms of loss or gain of electrons
	Metals lose electrons from their valence shell to form a positively charged cation (oxidation = loss of electrons) Non-metals gain electrons from their valence shell to form a negatively charged anion (reduction = gain of electrons)
3.3	be able to draw dot-and-cross diagrams to show electrons in cations and anions
	
3.4	be able to describe ionic crystals as giant lattices of ions
	Ions form crystalline giant lattice structures in which the ions are arranged in a regular pattern so the positive charges cancel out the negative charges
3.5	know that ionic bonding is the result of strong net electrostatic attraction between ions
	An **ionic bond** is the strong net electrostatic force of attraction between oppositely charged ions (in which the attraction between the ions acts in all directions)
3.6	understand the effects of ionic radius and ionic charge on the strength of ionic bonding
	Ionic radius is the measure of the size of an ion - Ionic radii increase with increasing negative charge - Ionic radii decrease with increasing positive charge - The strength of ionic bonding depends on: 1. Size of Ions 2. Charge of Ions In general, the smaller the ions and the larger the charge on the ions, (the higher the charge density), the stronger the ionic bonding

3.7	understand reasons for the trends in ionic radii down a group in the Periodic Table, and for a set of isoelectronic ions, including N^{3-} to Al^{3+}
	As we go down the group, *the ionic radii increases* (as there are more shells) For isoelectronic ions, (which means having the same electronic configuration) *The ionic radius decreases as the number of protons increases (as the electrons are more strongly attracted to the nucleus and hence pulled closer to it*
3.8	understand the meaning of the term 'polarisation' as applied to ions
	Theoretical lattice energies assume a perfect ionic model in which the ions are 100% spherical and the attractions are purely electrostatic However, the measured lattice energy differs from the theoretical lattice energy as we also have to take into account the covalent character when there is polarization of the anions In an ionic lattice, the positive ion will attract the electrons of the anion, distorting the electron density of the anion This polarization of the anions creates some degree of sharing of electrons between the two nuclei (i.e. covalent bonding) **Polarisation** is *the distortion of the electron density (electron cloud) of an anion by a cation*
3.9	understand that the polarising power of a cation depends on its radius and charge, and the polarisability of an anion also depends on its radius and charge
	A CATION with a **high charge** and **small radius** has a large polarizing power as it has a higher charge density An ANION with a **high charge** and **large radius** will be polarized more easily
	Further suggested practical The migration of ions in a U-tube using copper(II) chromate solution or on a microscope slide using potassium manganate(VII) crystals

3B: Covalent bonding

Students will be assessed on their ability to:

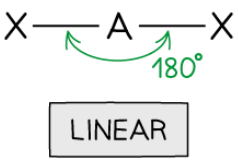
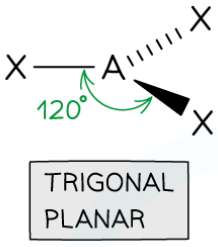
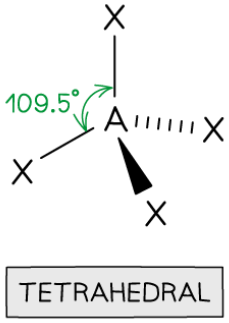
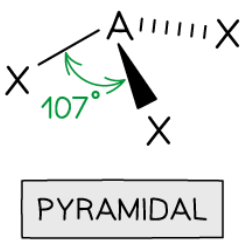
3.10	understand that *covalent bonding is the strong electrostatic attraction between two nuclei and the shared pair of electrons between them*, based on the evidence: i the physical properties of giant atomic structures ii electron density maps for simple molecules
3.11	be able to draw dot-and-cross diagrams to show electrons in covalent substances, including: i molecules with single, double and triple bonds ii species with dative covalent (coordinate) bonds, including Al_2Cl_6 and the ammonium ion
i	
ii	A **dative covalent bond** (a.k.a coordinate bond) is *the bond formed when an empty orbital of one atom overlap with an orbital containing a lone pair of electrons of another atom* !Note that the two electrons are shared and that both of them come from the same atom  Al_2Cl_6 (Aluminum Chloride)  Ammonium ion (NH_4^+)
3.12	be able to describe the different structures formed by giant lattices of carbon atoms, including graphite, diamond and graphene, and discuss the applications of each
	Carbon allotropes only contain carbon atoms but they have different physical properties due to their bonding arrangements such as Diamond, Graphite, Graphene and C_{60} Fullerene Giant covalent lattices such as Diamond, Graphite and Graphene have very high melting and boiling points, most are insoluble in water and most do not conduct electricity but there are exceptions such as Graphite! ***Graphite*** Each carbon atom is covalently bonded to three other carbon atoms, forming layers of hexagon and leaving one free electron per carbon atom It is due to these delocalised electrons that can move freely throughout the structure and carry charge when a potential difference is applied that graphite can conduct electricity It is soft and malleable as there are only weak intermolecular forces between the layers so the layers can slide over each other when a force is applied However! It has high melting and boiling points due to the strong covalent bonds

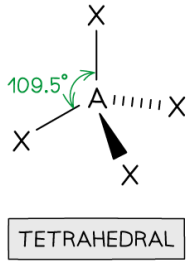
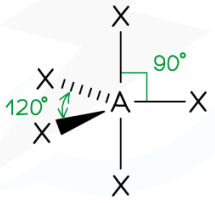
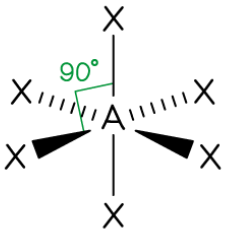
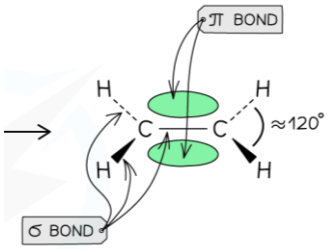
	between its *atoms within each layer* ***Diamond*** Each carbon atom is covalently bonded to four other carbon atoms, forming a tetrahedron (3-dimensional lattice structure) (and there are NO intermolecular forces, only strong covalent bonds making it hard and rigid) ***Graphene*** Basically, one layer of graphite makes it a graphene This means ITS ONLY ONE ATOM THICK!
3.13	understand the meaning of the term 'electronegativity' as applied to atoms in a covalent bond
	Electronegativity is the ability of an atom to attract a bonding pair of electrons in a covalent bond In general, electronegativity increases across a period → group a down decreases ↓ Most electronegative element 
3.14	know that ionic and covalent bonding are the extremes of a continuum of bonding type and be able to explain this in terms of electronegativity differences, leading to bond polarity in bonds and molecules, and to ionic bonding if the electronegativity is large enough
	Atoms of elements with large electronegativity differences (of 1.7 or more) tend to form ionic bonds (such as NaCl) Atoms of elements with similar electronegativity (less than 1.7) tend to form covalent bonds (such as HCl) Intermediate electronegativity differences between covalently bonded atoms lead to polarity in the bond When two atoms have the SAME electronegativity, it is non-polar A polar covalent bond is formed between two atoms where the bonding electrons are unequally distributed (because of this; one atom carries a slight negative charge and the other a slight positive charge) 
3.15	be able to distinguish between polar bonds and polar molecules and predict whether or not a given molecule is likely to be polar
	Some molecules have *polar bonds* but are overall not polar because the polar bonds in the molecule are arranged in such way that the individual *dipole moments cancel each other out* so molecule becomes *non-polar*
	Further suggested practical Determine the effect of an electrostatic force on jets of liquids (water, ethanol and cyclohexane) and use the results to determine whether the molecules are polar or non-polar

3C: Shapes of molecules

used to obtain basic shape

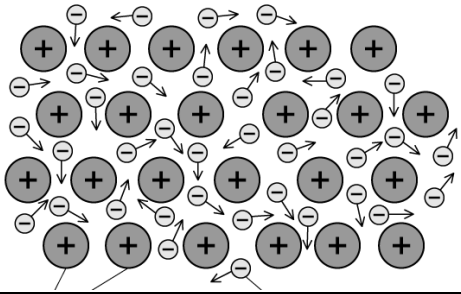
Students will be assessed on their ability to:

3.16	understand the principles of the electron-pair repulsion theory, used to interpret and predict the shapes of simple molecules and ions
used to estimate bond angles →	VSEPR (valence shell electron pair repulsion) theory (EPR for short) states that: • The shape of a molecule or ion is caused by repulsion between the pairs of electrons, both bonding pairs and lone pairs that surround the central atom • The electron pairs arrange themselves around the central atom into a position of maximum separation and minimum repulsion • Lone pair-lone pair repulsion > lone pair-bonding pair repulsion > bond pair-bond pair repulsion
3.17	understand the terms 'bond length' and 'bond angle'
	Bond length is the distance between the nuclei of two atoms in a covalent bond The shorter the bond length, the stronger the strength of the covalent bond Triple bonds are the shortest and strongest covalent bonds due to the large electron density between the nuclei of the two atoms C—C C=C C≡C 347 kJ/mol 614 kJ/mol 839 kJ/mol Bond angle is the angle between bonds :33
3.18	know and be able to explain the shapes of, and bond angles in, $BeCl_2$, BCl_3, CH_4, NH_3, NH_4^+, H_2O, CO_2, gaseous PCl_5, SF_6 and C_2H_4
	How to explain shapes?! 1. State the number of bonding pairs and lone pairs in that compound 2. State that the electron pairs arrange themselves around the central atom into a position of maximum separation and minimum repulsion 3. If there are lone pairs, state that lone pairs repel more than bonding pairs (which reduces bond angle) 4. If there are no lone pairs, state that the electron pairs repel equally 5. Finally, state the actual shape and bond angle
	$BeCl_2$ **2 bond pairs** 0 lone pairs **Linear** Bond angle is **180** 
	BCl_3 **3 bond pairs** 0 lone pairs **Trigonal planar** Bond angle is **120** 
	CH_4 **4 bond pairs** 0 lone pairs **Tetrahedral** Bond angle is **109.5** 
	NH_3 **3 bond pairs** 1 lone pair **Trigonal pyramidal** Bond angle is **107** Lp-lp repulsion is greater than bp-bp repulsion so angle is slightly less than 109.5 like in CH_4 

	NH_4^+ **4 bond pairs** 0 lone pairs **Tetrahedral** Bond angle is **109.5**  TETRAHEDRAL	H_2O **2 bond pairs** **2 lone pairs** **V-shaped** Bond angle is **104.5**  NON–LINEAR (BENT)
	CO_2 **2 bond pairs** 0 lone pairs **Linear** Bond angle is **180**  LINEAR	PCl_5 **5 bond pairs** 0 lone pairs **Trigonal bipyramidal** Bond angle is **90 and 120**  TRIGONAL BIPYRAMIDAL
	SF_6 **6 bond pairs** 0 lone pairs **Octahedral** Bond angle is **90 and 180**  OCTAHEDRAL	C_2H_4 **3 bond pairs** 0 lone pairs **Trigonal planar** Bond angle is **120**  ETHENE IS A PLANAR MOLECULE AS ALL THE BONDS LIE IN THE SAME PLANE
3.19	be able to apply the electron-pair repulsion theory to predict the shapes of, and bond angles in, molecules and ions analogous to those in 3.18	

3D: Metallic bonding

Students will be assessed on their ability to:

3.20	understand that metals consist of giant lattices of metal ions in a sea of delocalised electrons
	 Metals form giant metallic lattices surrounded by a sea of delocalized electrons
3.21	know that metallic bonding is the strong electrostatic attraction between metal ions and the delocalised electrons
	Metallic bonding is the strong electrostatic attraction between positively charged metal ions and the negatively charged delocalised electrons

3.22	be able to use the models in 3.20 and 3.21 to interpret simple properties of metals, including electrical conductivity and high melting temperature
	Metals have high melting and boiling points due to the strong electrostatic attraction between the ions and the delocalized electrons Metals can conduct electricity as the delocalized electrons can flow freely throughout the structure when a potential difference is applied Metals are malleable as the ions can slide past each other (while still being held together by the electrons) An alloy (a mixture of metal atoms) breaks up the regular lattice structure due to the different sizes of the metal ions, making it much stronger than pure metals